Unit 7
Chemical Equations and Reactions

What is a Chemical Equation?
A Chemical Equation is a written representation of the process that occurs in a chemical reaction. A chemical equation is written with the Reactants on the left side of an arrow and the Products of the chemical reaction on the right side of the equation. The head of the arrow typically points toward the right or toward the product side of the equation, although reactions may indicate equilibrium with the reaction proceeding in both directions simultaneously.

The elements in an equation are denoted using their symbols. **Coefficients** next to the symbols indicate the **stoichiometric** numbers. Subscripts are used to indicate the number of atoms of an element present in a chemical species.

An example of a chemical equation may be seen in the combustion of methane:

$$\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$$

Balancing Equations Notes
An equation for a chemical reaction in which the number of atoms for each element in the reaction and the total charge are the same for both the reactants and the products. In other words, the mass and the charge are balanced on both sides of the reaction.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>+</td>
<td>used to separate one reactant or product from another</td>
</tr>
<tr>
<td>→</td>
<td>used to separate the reactants from the products - it is pronounced &quot;yields&quot; or &quot;produces&quot; when the equation is read</td>
</tr>
<tr>
<td>⇌</td>
<td>used when the reaction can proceed in both directions - this is called an equilibrium arrow and will be used later in the course</td>
</tr>
<tr>
<td>(g)</td>
<td>indicates that the substance is in a gaseous state</td>
</tr>
<tr>
<td>↑</td>
<td>an alternative way of representing a substance in a gaseous state</td>
</tr>
<tr>
<td>(s)</td>
<td>indicates that the substance is in a solid state</td>
</tr>
<tr>
<td>↓</td>
<td>an alternative way of representing a substance in a solid state</td>
</tr>
<tr>
<td>(aq)</td>
<td>indicates that the substance is dissolved in water - the aq comes from aqueous</td>
</tr>
<tr>
<td>(l)</td>
<td>Identifies a phase state as pure liquid</td>
</tr>
<tr>
<td>△</td>
<td>indicates that heat is applied to make the reaction proceed</td>
</tr>
</tbody>
</table>
LAW OF CONSERVATION OF MASS

In all chemical equations the LAW OF CONSERVATION OF MASS must be met. Matter can NOT be created nor destroyed in a chemical reaction.

Remember, in a chemical reaction, the atoms/ions are simply rearranged to form new substances.

Therefore, chemical equations MUST be balanced.

WHAT IS A “BALANCED” CHEMICAL EQUATION?
A balanced chemical equation is one in which each side of the equation has the same number of atoms/ions of each element.

Example: \( \text{Al (s)} + \text{O}_2 (g) \rightarrow \text{Al}_2\text{O}_3 (s) \)

Not Balanced

\[
\begin{array}{cc}
1 & \text{Al} \\
2 & \text{O} \\
\end{array}
\quad \begin{array}{cc}
2 & \text{Al} \\
3 & \text{O} \\
\end{array}
\]

Balanced

\[
\begin{array}{cc}
4 & \text{Al (s)} \\
6 & \text{O} \\
\end{array}
\quad \begin{array}{cc}
3 & \text{O}_2 (g) \\
\quad \begin{array}{cc}
2 & \text{Al}_2\text{O}_3 (s) \\
4 & \text{Al} \\
6 & \text{O} \\
\end{array}
\end{array}
\]

RULES FOR BALANCING CHEMICAL EQUATIONS

1. Write the correct chemical formulas for all of the reactants and the products.

2. Write the formulas of the reactants on the LEFT of the reaction arrow; write the formulas of the products on the RIGHT of the reaction arrow.

3. COUNT the total number of atoms/ions of each element in the reactants and the total number of atoms/ions of each element in the products.

** A polyatomic ion that appears unchanged on both sides of the equation is counted as a single unit.

4. Balance the elements one at a time using coefficients.
   - A coefficient is a small WHOLE number that is written in front of a chemical formula in a chemical equation.
• When no coefficient is written, the coefficient is assumed to be 1.

• It is best to begin with elements OTHER THAN hydrogen and oxygen. These elements often occur more than twice in equations.

• ** You must NOT attempt to balance the equation by changing subscripts in chemical formulas!!!!!!

5. Check each atom/ion, or polyatomic ion to be sure that the equation is correctly balanced.

6. Finally, make sure that all of the coefficients are in the LOWEST possible whole number ratios. (At least one of the coefficients must be a prime number!)

Use coefficients to make sure the number of atoms is the same on both sides of the equation.

1. $\underline{2} \ \text{H}_2 + \underline{\text{O}_2} \rightarrow \underline{\text{2} \ \text{H}_2\text{O}}$

2. $\underline{2} \ \text{HCl} + \underline{\text{Zn}} \rightarrow \underline{\text{ZnCl}_2} + \underline{\text{H}_2}$

3. $\underline{2} \ \text{Al} + \underline{3} \ \text{CaS} \rightarrow \underline{\text{Al}_2\text{S}_3} + \underline{3} \ \text{Ca}$

4. Write the skeleton equation for the reaction of solid iron and gaseous chlorine react to produce a solid iron (III) chloride

   Write unbalanced equation
   $\text{Fe (s)} + \text{Cl}_2 (g) \rightarrow \text{FeCl}_3 (s)$

   Write balanced equation
   $2 \ \text{Fe (s)} + 3 \ \text{Cl}_2 (g) \rightarrow 2 \ \text{FeCl}_3 (s)$

Diatomic Elements

<table>
<thead>
<tr>
<th>Diatomic Elements are always diatomic (written with a subscribe of 2) when they are in their elemental form</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Hydrogen $\text{H}_2$ 5. Chlorine $\text{Cl}_2$</td>
</tr>
<tr>
<td>2. Nitrogen $\text{N}_2$ 6. Iodine $\text{I}_2$</td>
</tr>
<tr>
<td>3. Oxygen $\text{O}_2$ 7. Bromine $\text{Br}_2$</td>
</tr>
<tr>
<td>4. Fluorine $\text{F}_2$</td>
</tr>
</tbody>
</table>
Types of Chemical Reactions Notes

- **Synthesis** - two or more elements or compounds combine to form one compound.
- **Decomposition** - a single compound decomposes into two or more elements or smaller compounds.
- **Single Replacement** - a metal will replace a less active metal in an ionic compound OR a nonmetal will replace a less active nonmetal.
- **Double Replacement** - the metals in ionic compounds switch places.
- **Combustion** - an organic compound containing carbon, hydrogen and sometimes oxygen reacts with oxygen gas to form carbon dioxide and water.

**Examples**

\[ \text{Cu(s)} + \text{AgNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{Ag(s)} \]

Single Replacement

\[ \text{Fe(s)} + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) \]

Combustion

\[ \text{CuCl}_2(aq) + \text{Pb(NO}_3)_2(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{PbCl}_2(s) \]

Double Replacement

\[ \text{CaCO}_3(s) \rightarrow \text{CaO(s)} + \text{CO}_2(g) \]

Decomposition

---

**Synthesis**

Definition - two or more substances react to form 1 product. Usually releases energy, *EXOTHERMIC*. Combination reactions that contain oxygen as a reactant can also be considered combustion.

\[ A + X \rightarrow AX \]

- \( 4 \text{ Fe} (s) + 3 \text{ O}_2 (g) \rightarrow 2 \text{ Fe}_2\text{O}_3 (s) \)
- \( \text{CaO} (s) + \text{H}_2\text{O} (l) \rightarrow \text{Ca(OH)}_2 (s) \)

One example of a synthesis reaction is the combination of iron and sulfur to form iron (II) sulfide:

\[ 8 \text{ Fe} + \text{S}_8 \rightarrow 8 \text{ FeS} \]

**DECOMPOSITION**

Definition - A single compound breaks down into 2 or more elements or compounds

\[ AX \rightarrow A + X \]

- \( 2\text{NaN}_3(s) \rightarrow 2\text{Na}(s) + 3\text{N}_2 (g) \)
- \( 2\text{KClO}_3 (s) \rightarrow 2\text{KCl} (s) + 3\text{O}_2 (g) \)
CaCO₃ (s) → CaO (s) + CO₂ (g)

***These reactions often require an energy source as an initiator. Energy sources can be heat, light, or electricity. They are usually **ENDOTHERMIC**, **.

One example of a decomposition reaction is the electrolysis of water to make oxygen and hydrogen gas:

2H₂O ---> 2 H₂ + O₂

- **Combustion**: Definition - Oxygen gas combines with a substance and releases energy in the form of light or heat. So combustion reactions are usually *exothermic*. Combination reactions that contain oxygen as a reactant can also be considered combustion.

        A + O₂ →
        C(s) + O₂(g) → CO₂(g) + energy
        4 Fe (s) + 3O₂ (g) → 2Fe₂O₃ (s) + energy

*For hydrocarbons:*

        CₓHᵧ + [x + (y/4)] O₂ → xCO₂ + (y/2)H₂O
        C₃H₈(g) + O₂(g) → CO₂ (g) + H₂O (g) + light + heat

- **Single Replacement**: Definition - one ion replaces another in a compound.

        AB + C --> AC + B

    One example of a single displacement reaction is when magnesium replaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

        Mg + 2 H₂O ---> Mg(OH)₂ + H₂

- **Double Replacement**: Definition - two ions replace each other or switch places in compounds.

        AB + CD --> AC + BD

    One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

        Pb(NO₃)₂ + 2 KI ---> PbI₂ + 2 KNO₃
Sample Problems (the solutions are in the next section)

<table>
<thead>
<tr>
<th>List the type of the following reactions.</th>
<th>Solutions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) NaOH + KNO₃ → NaNO₃ + KOH</td>
<td>1) double replacement</td>
</tr>
<tr>
<td>2) CH₄ + 2 O₂ → CO₂ + 2 H₂O</td>
<td>2) combustion</td>
</tr>
<tr>
<td>3) 2 Fe + 6 NaBr → 2 FeBr₃ + 6 Na</td>
<td>3) single replacement</td>
</tr>
<tr>
<td>4) CaSO₄ + Mg(OH)₂ → Ca(OH)₂ + MgSO₄</td>
<td>4) double replacement</td>
</tr>
<tr>
<td>5) NH₄OH + HBr → H₂O + NH₄Br</td>
<td>5) acid-base (double replacement)</td>
</tr>
<tr>
<td>6) Pb + O₂ → PbO₂</td>
<td>6) synthesis</td>
</tr>
<tr>
<td>7) Na₂CO₃ → Na₂O + CO₂</td>
<td>7) decomposition</td>
</tr>
</tbody>
</table>

Practice

Balancing Chemical Equations Worksheet

Balance the following chemical equations using coefficients

1. 1Al(OH)₃(s) + 3HCl (aq) → 1AlCl₃ (aq) + 3H₂O (l)

2. 3Fe₂O₃ (s) + 1CO (g) → 2Fe₃O₄(s) + 1CO₂ (g)

3. 4FeO (s) + 1O₂ (g) → 2Fe₂O₃ (s)

4. 2C₆H₆ (l) + 15O₂ (g) → 12CO₂ (g) + 6H₂O (g)

5. 3Ca(OH)₂ (aq) + 2H₃PO₄ (aq) → 6H₂O (l) + 1Ca₃(PO₄)₂ (s)

6. 2I₄O₉ (s) → 1I₂O₆(s) + 3I₂ (s) + 6O₂ (g)
   **there’s another way to balance this equation…can you figure it out?**

7. 2Eu (s) + 6HF (g) → 2EuF₃ (s) + 3H₂ (g)

8. 3NaHCO₃ (aq) + 1C₆H₈O₇ (aq) → 3CO₂ (g) + 3H₂O (l) + 1Na₃C₆H₅O₇ (aq)
9. \( \text{Ni (s)} + 4\text{CO (g)} \rightarrow \text{Ni(CO)}_4 (g) \)

10. \( \text{K}_2\text{PtCl}_4 (aq) + 2\text{NH}_3 (aq) \rightarrow \text{Pt(NH}_3)_2\text{Cl}_2 (s) + 2\text{KCl (aq)} \)

**Write the following chemical equations and balance using coefficients.**

1. Liquid mercury reacts with liquid bromine to produce solid mercury (II) Bromide

   \[
   \text{Hg (l)} + \text{Br}_2 (l) \rightarrow \text{HgBr}_2 (s) \text{ balanced}
   \]

2. Solid calcium carbonate decomposes upon heating to produce solid calcium oxide and carbon dioxide gas.

   \[
   \text{CaCO}_3 (s) \rightarrow \text{CaO (s)} + \text{CO}_2 (g) \text{ balanced}
   \]

3. Solid calcium will react with liquid water to produce aqueous calcium hydroxide and hydrogen gas.

   \[
   \text{Ca (s)} + 2\text{H}_2\text{O (l)} \rightarrow \text{Ca(OH)}_2 (aq) + \text{H}_2 (g)
   \]

4. Butane gas (\( \text{C}_4\text{H}_{10} \)) will react with oxygen gas to produce carbon dioxide gas and water vapor.

   \[
   2 \text{C}_4\text{H}_{10} (g) + 13 \text{O}_2 (g) \rightarrow 8 \text{CO}_2 (g) + 10 \text{H}_2\text{O (g)}
   \]

5. Solid aluminum will react with oxygen gas to produce solid aluminum oxide.

   \[
   4\text{Al (s)} + 3\text{O}_2 (g) \rightarrow 2\text{Al}_2\text{O}_3 (s)
   \]

6. Aluminum metal is oxidized by oxygen (from the air) to form aluminum oxide.

   \[
   4 \text{Al (s)} + 3\text{O}_2 \rightarrow 2 \text{Al}_2\text{O}_3
   \]

7. Sodium oxide reacts with carbon dioxide to form sodium carbonate.

   \[
   \text{Na}_2\text{O} + \text{CO}_2 \rightarrow \text{Na}_2 \text{CO}_3
   \]

8. Calcium metal reacts with water to form calcium hydroxide and hydrogen gas.

   \[
   \text{Ca (s)} + 2\text{H}_2\text{O} \rightarrow \text{Ca (OH)}_2 + \text{H}_2 (g)
   \]

9. Potassium nitrate decomposes to form potassium nitrite and oxygen.

   \[
   2 \text{KNO}_3 (s) \rightarrow 2 \text{KNO}_2 (s) + \text{O}_2 (g)
   \]

10. Barium metal reacts with Iron (III) sulfate to produce barium sulfate and iron metal.

    \[
    3 \text{Ba (s)} + \text{Fe}_2(\text{SO}_4)_3 \rightarrow 3 \text{BaSO}_4 + 2 \text{Fe (s)}
    \]
11. Barium chloride reacts with sodium sulfate to produce barium sulfate and sodium chloride.

\[ \text{BaCl}_2 \text{ (aq)} + \text{Na}_2\text{SO}_4 \text{ (aq)} \rightarrow \text{BaSO}_4 \text{ (s)} + 2 \text{NaCl} \text{ (aq)} \]

Types of Chemical Reactions

*Directions*
(a) Write and balance the given equation.
(b) Indicate the type of chemical reaction represented.

1. Iron reacts with oxygen gas to produce Iron (III) oxide.
   (a) \(4 \text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3\)
   (b) Synthesis

2. Propane \((\text{C}_3\text{H}_8)\) reacts with oxygen gas to produce carbon dioxide and water.
   (a) \(\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 4\text{H}_2\text{O} + 3\text{CO}_2\)
   (b) Combustion

3. Bromine gas reacts with potassium iodide to produce potassium bromide and iodine gas.
   (a) \(\text{Br}_2 \text{ (g)} + 2\text{KI} \rightarrow 2\text{KBr} + \text{I}_2\)
   (b) Single displacement

4. Hydrogen peroxide will produce water and oxygen gas if left in sunlight.
   (a) \(2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2\)
   (b) Decomposition

5. White Phosphorous reacts with oxygen gas to produce tetraphosphorous decoxide.
   (a) \(4 \text{P} + 5\text{O}_2 \rightarrow \text{P}_4\text{O}_{10}\)
   (b) Synthesis

6. Iron (III) Chloride reacts with sodium hydroxide to produce Iron (III) hydroxide and sodium chloride.
   (a) \(\text{FeCl}_3 + 3\text{NaOH} \rightarrow \text{Fe(OH)}_3 + 3\text{NaCl}\)
   (b) Double Replacement

7. Iron (III) oxide reacts with hydrogen gas to produce iron and water.
   (a) \(\text{Fe}_2\text{O}_3 + 3\text{H}_2 \rightarrow 2\text{Fe} + 3\text{H}_2\text{O}\)
   (b) Single Replacement

8. Octane \((\text{C}_8\text{H}_{18})\) reacts with oxygen gas to produce carbon dioxide and water.
   (a) \(2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 18\text{H}_2\text{O} + 16\text{CO}_2\)
   (b) Combustion
9. Calcium carbonate reacts with aluminum phosphate to produce calcium phosphate and aluminum carbonate.

(a) \[ 3 \text{CaCO}_3 + 2 \text{AlPO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + \text{Al}_2(\text{CO}_3)_3 \]
(b) Double Replacement

10. Aluminum hydroxide decomposes to produce aluminum oxide and water.

(a) \[ 2 \text{Al(OH)}_3 \rightarrow \text{Al}_2\text{O}_3 + 3 \text{H}_2\text{O} \]
(b) Decomposition

11. Zinc reacts with silver nitrate to produce zinc nitrate and silver.

(a) \[ \text{Zn} + 2 \text{AgNO}_3 \rightarrow \text{Zn(NO}_3)_2 + 2 \text{Ag} \]
(b) Single Replacement

12. Glucose \((\text{C}_6\text{H}_12\text{O}_6)\) reacts with oxygen gas to produce carbon dioxide and water.

(a) \[ \text{C}_6\text{H}_12\text{O}_6 + 6 \text{O}_2 \rightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O} \]
(b) Combustion

13. Potassium oxide reacts with water to produce potassium hydroxide.

(a) \[ \text{K}_2\text{O} + \text{H}_2\text{O} \rightarrow 2 \text{KOH} \]
(b) Synthesis

14. Lead (IV) oxide decomposes into lead (II) oxide and oxygen gas.

(a) \[ 2\text{PbO}_2 \rightarrow 2\text{PbO} + \text{O}_2 \]
(b) Decomposition

15. Hydrochloric acid (hydrogen chloride) reacts with barium hydroxide to produce water and barium chloride.

(a) \[ 2 \text{HCl} + \text{Ba(OH)}_2 \rightarrow 2 \text{H}_2\text{O} + \text{BaCl}_2 \]
(b) Double Replacement
More Practice
Change the coefficients to make the number of atoms of each element equal on both sides of the equation

1. Calcium metal reacts with water to form solid calcium hydroxide and hydrogen gas.
   \[ \text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 (s) + \text{H}_2(g) \]

2. Zinc hydroxide solution reacts with lithium to form lithium hydroxide solution and zinc metal.
   \[ \text{Zn(OH)}_2 (aq) + 2 \text{Li} \rightarrow 2 \text{LiOH} (aq) + \text{Zn} \]

3. Liquid propanol (C₃H₇OH) reacts with oxygen gas to form carbon dioxide gas and water vapor.
   \[ 2 \text{C}_3\text{H}_7\text{OH} (l) + 9 \text{O}_2 \rightarrow 6\text{CO}_2 (g) + 8 \text{H}_2\text{O} \]
   "*note that C₃H₇OH is a Liquid (l)"

4. Aluminum metal reacts with oxygen gas to form solid aluminum oxide.
   \[ 4 \text{Al} + 3 \text{O}_2 (g) \rightarrow 2 \text{Al}_2\text{O}_3 (s) \]

5. Liquid carbonic acid (hydrogen carbonate) decomposes into carbon dioxide gas and water.
   \[ \text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 (g) + \text{H}_2\text{O} \]

6. Lead (II) nitrate solution reacts with iron (III) chloride solution to form solid lead (II) chloride and Iron (III) nitrate solution.
   \[ 3 \text{Pb(NO}_3)_2 (aq) + 2 \text{FeCl}_3 (aq) \rightarrow 3 \text{PbCl}_2 (aq) + 2 \text{Fe(NO}_3)_3 (aq) \]

7. Aluminum metal reacts with silver sulfate solution to form aluminum sulfate solution and silver metal.
   \[ 2 \text{Al} + 3 \text{Ag SO}_4 (aq) \rightarrow \text{Al}_2(\text{SO}_4)_3 (aq) + 3 \text{Ag} \]

8. Methane gas (CH₄) reacts with oxygen gas to form carbon dioxide gas and water vapor.
   \[ \text{CH}_4 + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (g) \]

9. Iron metal reacts with bromine gas to form iron (III) bromide solid.
   \[ 2 \text{Fe} + 3 \text{Br}_2 (g) \rightarrow 2 \text{Fe Br}_3 \]

    \[ 2 \text{H}_2\text{O}_2 (aq) \rightarrow 2 \text{H}_2\text{O} + \text{O}_2(g) \]
Rules for Predicting Products of Chemical Reactions

1. Hydrocarbon + O₂ → CO₂ + H₂O (Combustion)
   a. 2C₄H₁₀ + 13 O₂ → 8 CO₂ + 10 H₂O

2. Metal Carbonate → Metal Oxide + CO₂ (Decomposition)
   a. MgCO₃ → MgO + CO₂
   b. Synthesis: Metal Oxide + CO₂ → Metal Carbonate

3. Metal Sulfites → Metal Oxide + SO₂ (Decomposition)
   a. CaSO₃ → CaO + SO₂
   b. Synthesis: Metal Oxide + SO₂ → Metal Sulfite

4. Metal Hydride + H₂O → Metal Hydroxide + H₂ (Double Replacement)
   a. KH + H₂O → KOH + H₂

5. Metal + H₂O → Metal Hydroxide + H₂ (Single Replacement)
   a. 2Na + 2H₂O → 2NaOH + H₂

6. Metal Oxide + H₂O → Metal Hydroxide (Synthesis)
   a. MgO + H₂O → Mg(OH)₂

7. Non-metal oxide + H₂O → ternary acid (Synthesis)
   a. N₂O₃ + H₂O → 2 HNO₂
   b. N₂O₅ + H₂O → 2 HNO₃
   c. CO₂ + H₂O → H₂CO₃
Predicting Products: Write the COMPLETE balanced equation

1. Hydrochloric acid (HCl) reacts with sodium hydroxide.
   \[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \] (remember HOH is also H\(_2\)O)

2. Sodium reacts with oxygen
   \[ 4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O} \]

3. Mercury (II) oxide \( \rightarrow \)
   \[ 2\text{HgO} \rightarrow 2\text{Hg} + \text{O}_2 \]

4. Zinc reacts with lead (II) Nitrate
   \[ \text{Zn} + \text{Pb(NO}_3\text{)}_2 \rightarrow \text{Pb} + \text{Zn(NO}_3\text{)}_2 \]

5. Silver nitrate reacts with calcium chloride
   \[ 2\text{AgNO}_3 + \text{CaCl}_2 \rightarrow 2\text{AgCl} + \text{Ca(NO}_3\text{)}_2 \]

6. C\(_7\)H\(_{16}\) reacts with oxygen
   \[ \text{C}_7\text{H}_{16} + 11\text{O}_2 \rightarrow 8\text{H}_2\text{O} + 7\text{CO}_2 \]

7. CH\(_3\)OH reacts with oxygen
   \[ 2\text{CH}_3\text{OH} + 2\text{O}_2 \rightarrow 4\text{H}_2\text{O} + 2\text{CO}_2 \]

8. Magnesium reacts with Fluorine
   \[ \text{Mg} + \text{F}_2 \rightarrow \text{MgF}_2 \]

9. Copper (II) Chloride \( \rightarrow \)
   \[ \text{CuCl}_2 \rightarrow \text{Cu} + \text{Cl}_2 \]

10. Aluminum reacts with Calcium Sulfide
    \[ 2\text{Al} + 3\text{CaS} \rightarrow \text{Al}_2\text{S}_3 + 3\text{Ca} \]

11. Potassium Hydroxide reacts with Zinc Chloride
    \[ 2\text{KOH} + \text{ZnCl}_2 \rightarrow 2\text{KCl} + \text{Zn(OH)}_2 \]

12. C\(_2\)H\(_2\) reacts with oxygen
    \[ 2\text{C}_2\text{H}_2 + 5\text{O}_2 \rightarrow 2\text{H}_2\text{O} + 4\text{CO}_2 \]

13. Sodium Iodide reacts with chlorine
    \[ 2\text{NaI} + \text{Cl}_2 \rightarrow 2\text{NaCl} + \text{I}_2 \]

14. Aluminum reacts with sulfur
    \[ 2\text{Al} + 3\text{S} \rightarrow \text{Al}_2\text{S}_3 \]
Balance the following equations:

1. \(6 \text{HCl} + 2 \text{Al} \rightarrow 2 \text{AlCl}_3 + 3 \text{H}_2\)
2. \(4 \text{HNO}_3 + \text{Zn} \rightarrow \text{Zn(NO}_3)_2 + 2 \text{NO}_2 + 2 \text{H}_2\text{O}\)
3. \(4 \text{HNO}_3 + \text{Sn} \rightarrow \text{SnO}_2 + 4 \text{NO}_2 + 2 \text{H}_2\text{O}\)

Write the balanced equation for the following chemical reaction AND the type of reaction that has occurred.

4. Sodium metal is added to sulfuric acid and produce sodium sulfate and hydrogen.
   \(2 \text{Na} (s) + \text{H}_2\text{SO}_4 (aq) \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2 (g)\)  Single Replacement

5. White phosphorus (P₄) reacts with chlorine to make phosphorus trichloride
   \(\text{P}_4 (s) + 6 \text{Cl}_2 (g) \rightarrow 4 \text{PCl}_3 (g)\)  Synthesis

6. Magnesium chlorate, is heated strongly until it decomposes into magnesium chloride and oxygen gas.
   \(\text{Mg(ClO}_3)_2 (s) \rightarrow \text{MgCl}_2 (s) + 3 \text{O}_2 (g)\)  Decomposition

Write the balanced equation for the product AND the type of reaction that has occurred.

7. Butane gas (C₄H₁₀) is burned completely in air.
   \(2 \text{C}_4\text{H}_{10} (g) + 13 \text{O}_2 (g) \rightarrow 8 \text{CO}_2 (g) + 10 \text{H}_2\text{O} (g)\)  Combustion

8. Iron metal is added to bromine
   \(2 \text{Fe} (s) + 3 \text{Br}_2 (l) \rightarrow 2 \text{FeBr}_3 (s)\)  Synthesis

9. Calcium metal is added to phosphoric acid.
   \(3 \text{Ca} (s) + 2 \text{H}_3\text{PO}_4 (aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2 (s) + 3 \text{H}_2 (g)\)  Single Replacement

10. Aluminum metal is added to a solution of iron (III) chloride.
    \(\text{Al(s)} + \text{FeCl}_3 (aq) \rightarrow \text{AlCl}_3(aq) + \text{Fe(s)}\)  Single Replacement

11. A piece of iron metal is exposed to oxygen gas (Fe product would form).
    \(4 \text{Fe} (s) + 3 \text{O}_2 (g) \rightarrow 2 \text{Fe}_2\text{O}_3(s)\)  Synthesis

12. Calcium metal is added to water.
    \(\text{Ca} (s) + 2 \text{H}_2\text{O} (l) \rightarrow \text{Ca(OH)}_2 (s) + \text{H}_2 (g)\)  Single Replacement

13. Potassium metal reacts with chlorine gas.
    \(2 \text{K} (s) + \text{Cl}_2 (g) \rightarrow 2 \text{KCl} (s)\)  Synthesis